THE PERIODIC TABLE

1. Periodicity of physical properties of period 3 elements

1.1 Atomic radius

- Across period three, atomic radius decreases from sodium to chlorine.
- Across the period, atomic number increases so the nucleus has a higher positive charge.
- This draws electrons slightly closer to the nucleus meaning the atomic radius decreases.

1.2 Ionic radius

- Across period 3, the ionic radius decreases from sodium to aluminium (positive ions) then increases from silicon to chlorine (negative ions).
- The positive ions have a decreasing ionic radius because, although the ions have the same electron configuration, the number of protons in the nucleus increases so nuclear attraction increases.
- The negative ions have an increasing ionic radius because the ions have gained electrons meaning there are now more electrons than protons. As a result, the nuclear attraction to the electrons is weaker so they are not pulled in as strongly.

1.3 Melting point

Sodium, magnesium and aluminium – melting point increases.

- They are giant metallic structures with metallic bonding.
- The bond strength increases because the charge of the metal ion increases and the atomic radius decreases.

Silicon - has the highest melting point of all the elements in period 3.

- Silicon is a macromolecule with strong covalent bonds linking all its atoms together

Phosphorus (P_4), sulfur (S_8) and chlorine (Cl_2) - their melting points are dependent on the strength of their van der Waals/ intermolecular forces.

- They are all simple covalent molecules.
- the more atoms in a molecule mean stronger van der Waals forces so sulfur has the greatest melting point of these three molecules. Chlorine will have the lowest melting point since it is only made up of two atoms.

Argon will have the lowest melting point of period 3 because it exists as single atoms so there are very weak van der Waals intermolecular forces which are easy to overcome.

1.4 Electrical conductivity

- Conductivity of a compound relies on the presence of charged particles (such as ions or electrons) which are free to move.
- Metallic compounds can conduct electricity due to the sea of delocalised electrons. Conductivity increases from sodium to aluminium because the number of delocalised electrons increases so more electrons are free to carry charge.
- The molecules silicon to chlorine are covalent compounds with no charged particles so they are non-conductors.

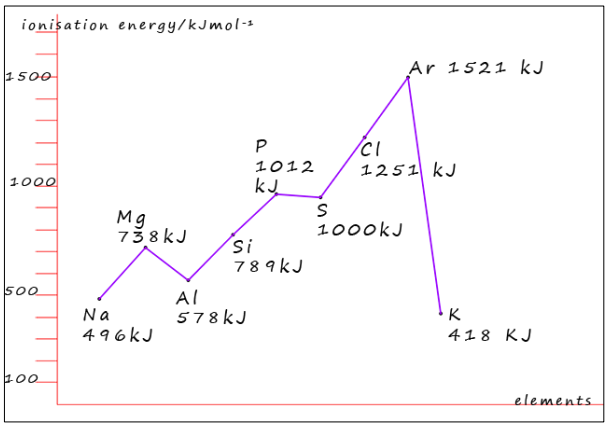
1.5 First ionisation energy

- The first ionisation energy is the energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous ions.
- First ionisation energy generally increases across period 3 - the number of electrons increases across the period while electron shielding remains the same.

Exceptions to the trend:

- The first ionisation energy of aluminium is lower than magnesium - electron is being removed from a 3p orbital rather than a 3s orbital. This means there is more electron shielding from the 3s orbital so the electron can be removed more easily.

- The first ionisation energy of sulfur is lower than phosphorus - sulfur is the first element to have an electron pair in a 3p orbital so an electron is being removed from a pair. The electrons in the pair repel each other slightly making it easier to remove an electron.



1.6 Ceramic

- Ceramics can either be giant covalent structures or giant ionic structures like magnesium oxide, aluminium oxide and silicon dioxide.
- The strong covalent or ionic bonds that hold the atoms / ions together in a giant ceramic structure make it strong, but also brittle.
- Have high melting point and are good electrical insulators

2. Periodicity of chemical properties of period 3 elements

2.1 Reactions with oxygen

- Sodium burns in oxygen with an orange flame to produce a white solid. $4Na + O_2 \rightarrow 2Na_2O$
- Magnesium burns in oxygen with a bright white flame to produce a white solid. $2Mg + O_2 \rightarrow 2MgO$
- Aluminium burns in oxygen to produce a white solid. $4Al + 3O_2 \rightarrow 2Al_2O_3$
- Silicon will only burn in oxygen if it is heated strongly enough. $Si + O_2 \rightarrow SiO_2$
- White phosphorus reacts spontaneously in the air to produce a white solid. $P_4 + 5O_2 \rightarrow P_4O_{10}$
- Sulfur burns with a blue flame in oxygen when heated gently. It produces a colourless gas of sulfur dioxide $S + O_2 \rightarrow SO_2$ which reacts further with oxygen to form sulfur trioxide. $2SO_2 + O_2 \rightarrow 2SO_3$
- Chlorine does not react with oxygen (directly).

2.2 Reactions with chlorine

- Sodium burns in chlorine with a bright orange flame to form white solid. $2Na + Cl_2 \rightarrow 2NaCl$
- Magnesium burns in chlorine with a bright white flame to form white solid. $Mg + Cl_2 \rightarrow MgCl_2$
- Aluminium burns in a stream of chlorine to produce pale yellow solid. $2Al + 3Cl_2 \rightarrow 2AlCl_3$. At high temperatures, aluminium chloride ($AlCl_3$) is converted into its molecular form Al_2Cl_6.
- If chlorine is passed over silicon powder, a colourless liquid forms. $Si + 2Cl_2 \rightarrow SiCl_4$
- White phosphorus burns spontaneously in chlorine to produce off-white solid. $P_4 + 10Cl_2 \rightarrow 4PCl_5$

2.3 Reactions with water

- Sodium undergoes a very exothermic reaction with cold water, producing hydrogen as bubbles and a colourless solution of sodium hydroxide. $2Na + 2H_2O \rightarrow 2NaOH + H_2$
- Magnesium reacts slowly with cold water, producing hydrogen. $Mg + H_2O \rightarrow Mg(OH)_2 + H_2$
- Magnesium burns in steam with a bright white flame to produce white magnesium oxide and hydrogen: $Mg + H_2O \rightarrow MgO + H_2$

2.4 Variation in oxidation numbers

Common oxidation states:

- +1 for group one atoms
- +2 for group two atoms
- -2 for oxygen (except when in a peroxide or F_2O)
- +1 for hydrogen (except in metal hydrides where it is -1)

- -1 for fluorine
- -1 for chlorine (except when bonded with fluorine or oxygen)

2.5 Period 3 oxides and water

● Sodium oxide is a strongly basic oxide which reacts exothermically with water to form a solution of sodium hydroxide (around pH 14). $Na_2O + H_2O \rightarrow 2NaOH$

● Magnesium oxide undergoes a slight reaction with water, forming some magnesium hydroxide ions. As these are only partially soluble, the pH of the resulting solution is about 9 since not many hydroxide ions are released into the solution. $MgO + H_2O \rightarrow Mg(OH)_2$

● Aluminium oxide does not react with water.

● Silicon dioxide does not react with water because it is difficult to break up the macromolecule.

● Phosphorus(V) oxide reacts violently with water to form phosphoric acid. $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$

● Sulfur dioxide reacts with water to form an acidic solution of sulfurous acid. $SO_2 + H_2O \rightarrow H_2SO_3$

● Sulfur trioxide reacts violently with water to produce sulfuric acid. $SO_3 + H_2O \rightarrow H_2SO_4$

2.6 Period 3 chlorides and water

● Sodium chloride and magnesium chloride are ionic compounds so they dissolve in water to form their ions.

● Silicon tetrachloride reacts with water in the air to produce hydrogen chloride. $SiCl_4 + 2H_2O \rightarrow SiO_2 + 4HCl$

● Phosphorus(V) chloride also reacts violently with water to produce hydrogen chloride. $PCl_5 + H_2O \rightarrow POCl_3 + 2HCl$

● Disulfur dichloride reacts slowly with water to produce a range of products. There are lots of examples of equations that could take place so it is unlikely this will be asked in an exam. Potential products include HCl, S and H_2S.

2.7 Chemical bonding in chlorides and oxides

In chlorides and oxides, the chemical bonding from sodium to aluminium is ionic.

In chlorides and oxides, the chemical bonding from silicon to sulfur is covalent.